

Ch 5 Gases

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Fun Chemistry/Physics

Many reactions we do yield gases they all behave in a similar fashion

S.I Pressure = $\frac{\text{Force}}{\text{area}}$ really $\frac{\text{total gas molecules colliding w/ container}}{\text{area}}$

How is pressure measured: Barometer or Manometer

Review units kPa, mmHg, torr, atm

or Physics in SI \rightarrow Pascal $\frac{\text{N}}{\text{m}^2}$

Gas Laws

Robert Boyle - Volume vs Pressure

describe his apparatus and graphs

Boyle's Law

$$VP = K \text{ and } V_1 P_1 = V_2 P_2$$

Charles's Law

Volume of gas increases linearly w/ T

Show graph to show how can be used to find Abs 0 b/c all gases extrapolate to 0 Volume @ 0 K

true direct Relationship.

Jacques Charles

1st Solo Balloon Flight

Now show

$$V = bT$$

\uparrow
constant

$$\text{but } V_1 T_2 = V_2 T_1$$

Avogadro's Law - equal V 's ^{of gases} ~~of particles~~ have same # of particles
@ const $T + P$

Thus $V = an$

Can use $V_1 n_2 = V_2 n_2$

all of these can be combined to give

$$V = R \left(\frac{Tn}{P} \right) \rightarrow PV = nRT$$

\uparrow
 k, b, a ideal gas constant

$0.08206 \frac{\text{Latm}}{\text{molK}}$

Derive
 $D = \frac{MP}{RT}$
 $M = \frac{mRT}{PV}$

other gas laws

Gay Lussac's $P_1 T_2 = P_2 T_1$

Combined $V_1 P_1 T_2 = V_2 P_2 T_1$

Important to know if 2 things are constant
the ideal gas law reduces to each of the other eqns

Gas Stoichiometry

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Avogadro's Law suggests There is a volume where 1 mole of a gas will be constant @ STP

use ideal gas law to solve. $22.42 \frac{L}{mol}$ @ STP

This does vary a little but remember ideal gases are hypothetical.

Dalton's Law of Partial Pressures

Dalton showed a mixture of gases ~~are~~ pressure all add together to give a total Pressure.

$$P_T = P_1 + P_2 + P_3 \dots$$

also in terms of mol fraction

$$X_i = \frac{n_i}{n_{total}}$$

$$X_i = \frac{P_i}{P_{total}}$$

lots of applications

$$b/c PV = nRT$$

Kinetic Molecular Theory

We have been throwing around "ideal" gas
and saying that an ideal gas is one that pertains to the
gas laws but they don't tell us why!

Why do gases behave according to ideal gas law?

Kinetic Molecular Theory was developed to explain this.

↳ attempts to explain ideal gas behavior

Show Slide w/ 4 statements

Explain how Real gases don't conform to these assumptions

Explain how V vs P , P vs T , V vs T , n + P , ~~other~~ Mixture
work

Don't Skip how ideal Gas Law derived →

Diffusion - mixing of gas ~~then~~ in absence of air currents

Effusion - passing of gas through a tiny orifice

Knowing ~~KE~~

$KE = \frac{1}{2}mv^2$ can derive

$$\frac{\text{Rate}_1}{\text{Rate}_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

AP Exam

$$\frac{\text{time}_1}{\text{time}_2} = \frac{\sqrt{M_1}}{\sqrt{M_2}}$$

$$\frac{\text{Dist}_1}{\text{Dist}_2} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Deriving the ideal Gas law

Appendix 2 derives via definition of velocity, momentum (pythagorean theorem)

Force $F = ma$ pressure $P = \frac{F}{a}$ collision freq.

Avogadro's #

to get

$$P = \frac{2}{3} \left[\frac{n N_A \left(\frac{1}{2} m \bar{U}^2 \right)}{V} \right]$$

P = pressure

n = moles

N_A = Avogadro's #

V = volume

m = particle mass

\bar{U}^2 = avg square velocity of each particle

Comes from pythagorean theorem

$N_A \frac{1}{2} m \bar{U}^2$ = Avg KE per a mole of gas particles

Now we get $P = \frac{2}{3} \left[\frac{n KE_{avg}}{V} \right]$ or $\frac{PV}{n} = \frac{2}{3} (KE_{avg})$

Since T is directly Prop to KE ← derived using KMT assumption

$\frac{PV}{n} = \frac{2}{3} KE_{avg} \propto T$

or $\boxed{\frac{PV}{n} \propto T}$
theory

$\frac{PV}{n} = RT$

Experiment

proves KMT

Temp proven to be about motion

$\boxed{\frac{PV}{n} = \frac{2}{3} KE_{avg} = RT}$

so $KE_{avg} = \frac{3}{2} RT$

↑ Kelvin temp is index of

Random gas motion in a substance

Root mean Square Velocity

\bar{U}^2 - average of squares of velocity of particles

take square root and get U_{rms}

$$U_{rms} = \sqrt{\bar{U}^2}$$

Derive U_{rms} by the following

$$(KE_{avg}) = N_A \left(\frac{1}{2} m \bar{U}^2 \right) \text{ and } KE_{avg} = \frac{3}{2} (RT)$$

$$N_A \left(\frac{1}{2} m \bar{U}^2 \right) = \frac{3}{2} (RT) \rightarrow \bar{U}^2 = \frac{3RT}{N_A m}$$

$$\sqrt{\bar{U}^2} = U_{rms} = \sqrt{\frac{3RT}{N_A m}}$$

m is in Kg
so $N_A m$ is mole of gas in
Kg M

$$U_{rms} = \sqrt{\frac{3RT}{M}}$$

universal gas const

Big Range of velocities in gases B/c particles

U_{rms} for O_2 @ STP is $500 \frac{m}{s}$ actual is graph 5.20

so a wide range and 5.21 shows that the ~~the~~ velocity ↑

w/ T

Real Gases

(3)

Ideal gases are hypothetical Concepts

No gas exactly follows ideal gas law. This behavior is approached by real gases under certain conditions

Let's examine how close the approximation of $k_B n T$ is

We'll use $PV = nRT$

Compare $\frac{PV}{nRT}$ vs P to see how Quantity of $\frac{PV}{nRT}$ changes w/ respect to P

For ideal gas $\frac{PV}{nRT}$ should be one @ all conditions

but graph shows only 1 @ low P

Then graph done for N_2 @ different T 's and again

@ higher T 's $\frac{PV}{nRT}$ approaches 1

So gases are ^{close to} ideal gases @ Low P + high T

Correcting for ~~ideal~~ Real Conditions

Vander waals forces

1910 Nobel Prize for work on Real Gases

Start w/ ideal gas law

$$P = \frac{nRT}{V}$$

assumes volumeless entities

but gases have finite volumes

allowing volume of real gas particles to take up space of container

to correct for this Vander waals said

$$V_{\text{container}} = V - nb$$

correction factor for V of molecules

$n = \# \text{ of moles}$

$b = \text{empirical constant}$

determined experimentally

So case by case basis

Now I get

$$P' = \frac{nRT}{(V - nb)}$$

Now let's factor in attractive forces of molecules

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IM attractions make P_{obs} smaller so - See fig 5.27

$$P_{obs} = (P' - \text{correction factor}) = \left(\frac{nRT}{V-nb} - \text{correction factor} \right)$$

Size of correction factor depends on $[gas] \propto \frac{n}{V}$

so higher $[]$ the lower the actual P

So correction factor

is $\propto \frac{n}{V}$ for large H 's of particles the H of interacting

particles depends on the square of $\frac{n}{V}$ or $\left(\frac{n}{V}\right)^2$

to explain show fig 5.28

each particle can interact in 2 variations. why squared.

N particles $N-1$ combinations and each pair count 2x

so $N(N-1)/2$ pairs and $N^2/2$
↑ Big sample so $N-1 \approx N$ ↑ why $[]^2$
B/c same for each particle

Thus $P_{obs} = P' - a \left(\frac{n}{V} \right)^2$

proportionality const, includes $\frac{1}{2}$ from $\frac{N^2}{2}$ and
is determined by observing gas behavior.

$$P_{\text{obs}} = \frac{nRT}{V - nb} - a \left(\frac{n}{V} \right)^2$$

obs P → P_{obs}
 V container → $V - nb$
 nb → $V_{\text{corrected}}$
 $a \left(\frac{n}{V} \right)^2$ → $P_{\text{corrected}}$

Rearranged to get

$$\underbrace{\left(P_{\text{obs}} + a \left(\frac{n}{V} \right)^2 \right)}_{P_{\text{ideal}}} \times \underbrace{V - nb}_{V_{\text{ideal}}} = nRT$$

See tables 5.3 for a+b values

a+b increase w/ size of molecule

But again @ high V the particle V's are negligible

high Temp particles move so fast they don't interact.

So IM forces can be shown here.

